**Solutions for Activity 2**

**Visual 1: Hydrogen Fuel Cell**

1. Write the equation for the overall cell reaction that occurs in a hydrogen fuel cell?

**Solution 2H2(g) + O2(g) 🡪 2H2O(l)**

1. Write down the half-reactions which occur at the cathode and anode in a hydrogen fuel cell?

(Use the relative strengths of oxidizing agents table provided)

**Solution**

**H2 (g)→2H+(s) + 2e- (at anode)**

**O2(g) + 4H+(s) + 4e- 🡪 2H2O(l) (at cathode)**

1. Calculate Eocell for a hydrogen fuel cell.

**Solution**

**Eocell = Eocathode – Eoanode = 1.23 – 0.00 = 1.23 V** (using the Standard Reduction Potentials table)

1. What is the source of oxygen and hydrogen for these fuel cells?

**Solution**  
**Oxygen is obtained from the atmosphere.**

**Only trace amounts of hydrogen gas are found in the atmosphere, therefore hydrogen for use in fuel cells must be extracted from other compounds. Hydrogen can be obtained from electrolysis of water or from hydrocarbons by a process called reforming.**

What does it make you wonder?

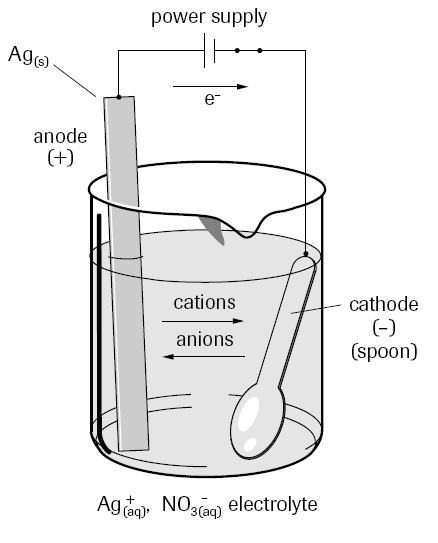
**The electrolysis of water is a very costly process and it requires a lot of energy. The process of reforming requires burning of fossil fuels which is environmentally-unfriendly.**

**Therefore, we need to find environmentally friendly ways such as nuclear, wind and solar energies to produce hydrogen if we want the wide-spread use of hydrogen fuel cells.**

**Visual 2: Electroplating**

1. Suppose you want to set up an electrolytic cell to electroplate a metal spoon with a thin layer of silver.
2. As part of the experimental design, draw the cell and label the electrodes, power supply, electrolyte, and the direction of electron and ion movement.

**Solution**



1. What variables must be considered when planning the electrolysis?

**Solution**

**Some variables that need to be considered when planning the electrolysis include: the selection, solubility, and concentration of the electrolyte; the potential difference that will be applied and the current to be used; the length time the cell will operate; the mass of silver. (Some of these variables are related to each other.)**

1. List some reasons for, and examples of, electroplating.

**Solution**

**Electroplating is used to coat a strong base metal with a surface that is more attractive, or corrosion resistant, or both. Some metals are plated with silver and gold for appearance, nickel for corrosion resistance, and chromium for both appearance and corrosion resistance.**

1. Electroplating solutions become dirty after few uses and need to be replaced. These electroplating solutions that can no longer be used are known as ‘spent’ electroplating solutions. How should the electroplating waste be managed?

**Solution**

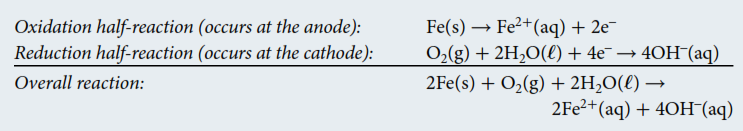
**‘Spent’ electroplating solutions contain some precious metal ions for example silver and gold which must be recovered. These ions can be recovered using various processes such as solvent extraction.**

**The resulting waste from these processes can be managed by identifying the type of waste for example if they are heavy metal waste (containing Cd, Cr, Pb or Hg), alkaline or acidic solutions, solvents and cyanide waste and treating them accordingly. These wastes that are potential hazards should be disposed of in a manner so that they do not contaminate nor damage our environment.**

What does it make you wonder?

**Are electroplating solutions discarded properly?**

**Visual 3: Rusting/Corrosion**



1. Determine the standard cell potential for the corrosion reaction.

**Solution**  
**Eocell = Eocathode – Eoanode = 0.40V - (-0.45V) = 0.85 V**

1. What is the electrolyte for the corrosion reaction?

**Solution**

**Carbonic acid formed from the reaction of rainwater with the carbon dioxide acts as the electrolyte for the corrosion process.**

1. Does the acid rain prevent or promote the rusting of iron? Why?

**Solution**

**Acid rain promotes the rusting of iron.  
The chemical compounds found in liquids like acid rain, seawater and the salt-loaded spray from snow-belt roads make them better electrolytes than pure water.**

1. What are some societal consequences of corrosion?

**Solution**  
**The corrosion of steel reinforcing bar in concrete can proceed out of sight and suddenly result in failure of a section of highway, the collapse of electrical towers, and damage to buildings, parking structure and bridges, etc., resulting in significant repair costs and endangering public safety.**

What does it make you wonder?

**Are there any ways to prevent rusting?**

**Visual 4: Corrosion Preventive Methods**

1. Does a protective layer of tin on iron provide effective protection against rusting?

**Solution**

**No, because tin is not a better oxidizing agent than iron. Since tin is less reactive than iron, tin acts as a cathode in each miniature galvanic cell on the surface of iron. Iron acts as the anode of each cell, which is its normal role when rusting.**

1. Explain why aluminum provides cathodic protection to an iron object?

**Solution**

**Aluminum provides cathodic protection to iron because it is more reactive than iron and acts as a sacrificial anode, thus preventing the iron from rusting.**

1. A zinc wire is connected to and buried with a pipeline when it is built. Why is this done? What are some of the environmental and safety issues associated with protecting and also not protecting pipelines.

**Solution**  
**A zinc wire is connected to protect the pipelines from corrosion.**

**Regular maintenance of pipelines by cathodic protection (as sacrificial anode is slowly destroyed by oxidation, it must be replaced periodically) is a costly process.**

**If pipelines are not protected against corrosion, failure of pipelines could result. Also rust could mix with water thus jeopardizing public health and safety.**

What does it make you wonder?

**Are there any ways in which rusting can be prevented permanently? Bridges and pipelines should be built with corrosion-resistant metals in the first place so that a lot of money is not wasted in their maintenance to protect them from corrosion.**

**Visual 5: Rechargeable Batteries**

1. Do rechargeable batteries act as galvanic cell or electrolytic cell?

**Solution  
Rechargeable batteries operate some of the time as galvanic cells and some of the time as electrolytic cells when they are recharged.**

1. The overall reaction for a lead-acid battery is:

Pb(s) + PbO2(s) + 4H+(aq) + SO4-2(aq) 🡪 2PbSO4(aq) + 2H2O(l)

1. Which material acts as the anode and which acts as the cathode in the reaction?

**Solution**  
**Pb acts as the anode and PbO2 acts as the cathode.**

1. Write down the balanced half-reactions.

**Solution**

**Pb(s) + SO42-(aq) 🡪 PbSO4(s) + 2e-**

**PbO2(s) + 4H+(aq)+ SO42-(aq) + 2e- 🡪 PbSO4(s) + 2H2O(l)**

1. Write the two half reactions and the overall cell reaction for the process when this battery is recharged.

**Solution**

**PbSO4(s) + 2e- 🡪 Pb(s) + SO42-(aq)**

**PbSO4(s) + 2H2O(l) 🡪 PbO2(s) + 4H+(aq) + SO42-(aq) + 2e-**

**2PbSO4(aq) + 2H2O(l) 🡪 Pb(s) + PbO2(s) + 4H+(aq) + 2SO42-(aq)**

1. What external voltage is required to recharge a lead-acid car battery?

**Solution**

**Eocell = Eocathode – Eoanode = -0.36 – (1.69) = -2.05 V**

1. After being discharged completely, even rechargeable batteries wear out. What are some of the environmental hazards and health risks caused by the improper use and disposal of lead-acid and NiCad batteries?

**Solution**

**One must be careful while handling lead-acid batteries as they contain sulphuric acid which is highly corrosive.**

**They must be disposed of properly because they contain lead and sulphuric acid, which are both hazardous to the environment and human health.**

**Discarded NiCad batteries release toxic cadmium. The toxicity of this substance makes it hazardous to the environment as cadmium can enter the food chain. Long-term exposure to low levels of cadmium can have serious medical effects on humans, such as high blood pressure and heart disease.**  
What does it make you wonder?

**In theory, worn out batteries should be recycled. In practice, however, many end up in landfills. What steps should be taken to ensure their proper disposal?**

**Visual 6 – Electrorefining of copper**

1) Write down the balanced half reactions which occur at cathode and anode.

**Cathode: Cu2+(aq) + 2e- 🡪 Cu(s)**

**Anode: Cu(s) 🡪 Cu2+(aq) + 2e-**

2) Why must the pure copper metal product form at the cathode?

**Solution  
 A metal product must form at the cathode during electrolysis because metal ions are positively charged, and must gain electrons (be reduced) to become atoms of metal.**

3) What is the minimum electric potential difference required for this cell?

**Solution**  
**The minimum electric potential difference required for this cell is theoretically zero**.

4) Why is it unlikely that your answer to question #2) is what is used? Discuss briefly.

**Solution  
The minimum potential difference is a theoretical minimum and would mean a very long reaction time. A higher voltage is used to get the reaction to occur rapidly. (Note that the choice of voltage is also influenced by the cost of electricity.)**

1. How do you think the anode mud formed in this process is treated?

**Solution  
 The anode mud is removed from the cell periodically and undergoes further processing to extract valuable materials.**

What do you want to know?

**Are there other metals which could be refined in the similar way?**

What does it make you wonder?

**Electrorefining of copper is essential for the manufacture of electrical wires.**

**What happens to the spent electrorefining solutions? How are they disposed off?**

**Visual 7 – Chlor-alkali process**

1) Why is the Chlor-alkali process commercially significant? What is the significance of the products produced in this process?

**Solution**

**Most of the chlorine and sodium hydroxide that are used commercially are produced by the chlor-alkali process. Sodium hydroxide and chlorine are two of the most extensively produced commercial chemicals.**

**Chorine is used to make laundry bleach, to bleach pulp for paper, to make compounds for treating water, as a disinfectant, and to make hydrochloric acid. A large portion of the chlorine that is produced every year goes into making polyvinyl chloride (PVC), a type of plastic.**

**Sodium hydroxide is used in making soaps and detergents, in the production of aluminum, and in the manufacturing of many different chemicals. It is also used in the pulp and paper industry to break down the lignin in woods that holds fibres together.**

2) Write down the balanced half reactions at the cathode and anode.

**Solution**

**At cathode: 2H20(l) + 2e- 🡪 2OH-(aq) + H2(g)**

**At anode: 2Cl-(aq) 🡪 Cl2(g) + 2e-**

3) What is the minimum electric potential difference required for this cell?

**Solution**

**Eocell = Eocathode – Eoanode = -0.83 – (1.36) = -2.19 V**

4) Chlorine is a controversial chemical. Although, chlorine and the products made from chlorine have been very beneficial to society, there are some concerns. What are some of the concerns that people have with the use of chlorine?

**Solution**

**Chlorine is very corrosive, and a very strong oxidizing agent. The chemical is harmful to humans by skin contact and especially harmful by breathing the vapour. Edema of the lungs and chronic bronchitis may result from exposure by inhalation.**

**Many groups are concerned with chlorine’s potential for harm to the environment and to human health are actively campaigning to have its use in water treatment and the pulp industry replaced by other chemicals. Some chlorine compounds are now banned. While these compounds initially seemed chemically inert, it has been discovered that they accelerate the breakdown of ozone in the stratosphere — “thinning” the ozone layer that absorbs some of our sun’s harmful UV radiation.**

**Visual 8 – Extraction of Reactive Metals**

1. Explain why sodium cannot be produced by electrolysis of sodium chloride solution in water.

**Solution**

**Sodium cannot be produced by electrolysis of sodium chloride solution because water is a stronger oxidizing agent than sodium ions therefore water will undergo reduction instead of sodium ions.**

1. Write the half reactions that occur at the cathode and anode in a Downs cell.

**Solution**

**Na+(l) + e- 🡪 Na(l) (at cathode)**

**2Cl-(l) 🡪 Cl2(g) + 2e- (at anode)**

1. Write the net ionic equation for the overall cell reaction.

**Solution**

**2Na+(l) + 2Cl-(l) 🡪 2Na(l) + Cl2(g)**

1. How can the industrial chemists ensure that only sodium ions (and no calcium ions) are reduced in the Downs cell?

**Solution**

**The potential difference (voltage) applied to the Downs cell must be controlled to reduce sodium ions but not calcium ions. (Note: sodium is more easily reduced than calcium)**

1. Calcium Chloride is added to reduce the melting point of sodium chloride in the Downs cell. Explain why calcium chloride does not affect the reduction of sodium ions.

**Solution**

**Calcium chloride does NOT affect the reduction of the sodium ions because Ca2+ is a weaker oxidant than Na+.**

1. Given that the melting point of sodium is only 98oC, is the sodium obtained through this cell solid or liquid?

**Solution**

**Liquid sodium will be obtained at the cathode**.

What does it make you wonder?

**Electrolysis has many industrial applications. Sodium extracted from the Downs cell is widely used for sodium vapour lamps.**